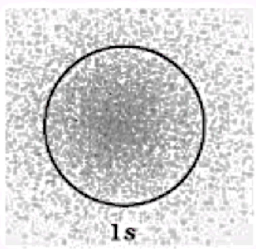
Assignment 7.3 Questions 20, 21, 62, 63, 67, 69, 73 – 81 odd

20) s = 2, p = 6, d = 10, f = 14 elements
Each block increases by 4, so the next two element blocks would hold 18 and 22 elements each.

21) Stability comes from the attraction between electrons and protons. Instability comes from the repulsion between electrons and other electrons.

As each electron is filled into an empty orbital, they have little repulsion (no other electrons in that orbital), but more nuclear attraction (there is also an additional proton in the nucleus.) Thus the atom is more stable. When the sublevel is half filled (e.g. N or Mn) the atom has more of these unpaired electrons. If one more electron is added, it must pair up with another electron. The repulsion between the electron pair is greater than the attraction created by the extra proton. Thus oxygen is less stable than nitrogen, and fluorine is even less stable.

62)



It is literally impossible to know exactly where an electron is at any one point in time. The act of looking causes it to move. An orbital is just the region where the electron is most likely to be. This cloud becomes more and more diffuse farther from the center, so we arbitrarily choose the region where the electron is 90% of the time.

63) 5p = 3 orbitals $3d_{z^2}$ = 1 orbital 4d = 5 orbitals
(n = 5) = 25 orbitals ($2n^2$ = 50 electrons = 25 orbitals)
(n = 4) = 16 orbitals ($2n^2$ = 32 electrons = 16 orbitals)

67)

Na	[↑↓] 1s	[↑↓] 2s	[↑↓ ↑↓ ↑↓] 2p	[↑] 3s	1 unpaired electron
Co	[Ar]	[↑↓] 4s	[↑↓ ↑↓ ↑ ↑ ↑] 3d		3 unpaired electrons
Kr	[Ar]	[↑↓] 4s	[↑↓ ↑↓ ↑↓ ↑↓ ↑↓] 3d	[↑↓ ↑↓ ↑↓] 4p	No unpaired electrons

69)
Si = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
Ga = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$
As = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$
Ge = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^2$
Al = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
Cd = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}$
S = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
Se = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$

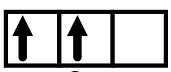
73) a. Iodine - $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^5$
(Indium also has 1 unpaired 5p electron, but metals don't form covalent bonds)
b. Benjaminium - [Rn] $7s^2\ 5f^{14}\ 6d^{10}\ 7p^6\ 8s^2$
c. Radon - [Xe] $6s^2\ 4f^{14}\ 5d^{10}\ 6p^6$
d. Mn - [Ar] $4s^2\ 3d^5$ (Cr is technically correct, but the reason why is beyond the scope of this course.)

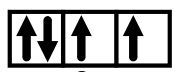
75) a. 18 electrons in 3rd energy level ($3s^2$, $3p^6$ and $3d^{10}$)
b. 30 electrons with d orbitals ($3d^{10}$, $4d^{10}$ and $5d^{10}$)
c. 8 electrons in p_z orbitals ($2p_z^2$, $3p_z^2$, $4p_z^2$, and $5p_z^2$)
d. 40 electrons have an "up" spin (half of the 80 total electrons)

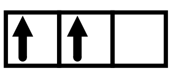
77) Skip (quantum numbers are no longer part of the AP curriculum)

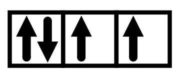
79) This configuration has ***no unpaired electrons***. Oxygen's ground state is $1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$. Having a configuration with $2p_y^2$ is less stable so it is an ***excited state***. When going back to the ground state, ***energy would be released***.

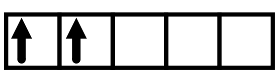
81) Elements with 2 unpaired ground state electrons:

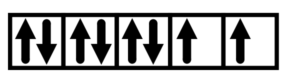
Carbon  2p

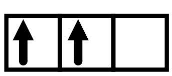
oxygen  2p

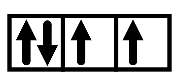
silicon  3p

sulfur  3p

titanium  3d

nickel  3d

germanium  4p

selenium  4p

You didn't have to draw orbital diagrams, but I wanted you to see why these all had 2 unpaired electrons.